Chemistry 129.01 - General Chemistry Workshop - Spring 2017

Week #3

Monday, February 6. Sessions 3 of the Greenhouse Gas Module.
Assigned reading: Sections 6.3-6.4

This week, we are going to refine our model of the structure of an atom. Last week looked at the Bohr model of the atom which proposes that there are only a few allowed states for electrons in "orbit" around a nucleus. Despite being successful at predicting the hydrogen atom line spectrum, the Bohr model fails to predict the spectrum of other atoms. We'll extend the ideas of Bohr a little bit and start the description of quantum mechanics. Start to think about the idea that matter (especially single electrons) acts somewhat like a wave. The mathematical description of these "waves" are often called "orbitals", similar to but not the same as a Bohr "orbit". In class, we'll focus on the shape of this orbitals and in ways to represent the orbital occupancy of the ground-state of atoms.

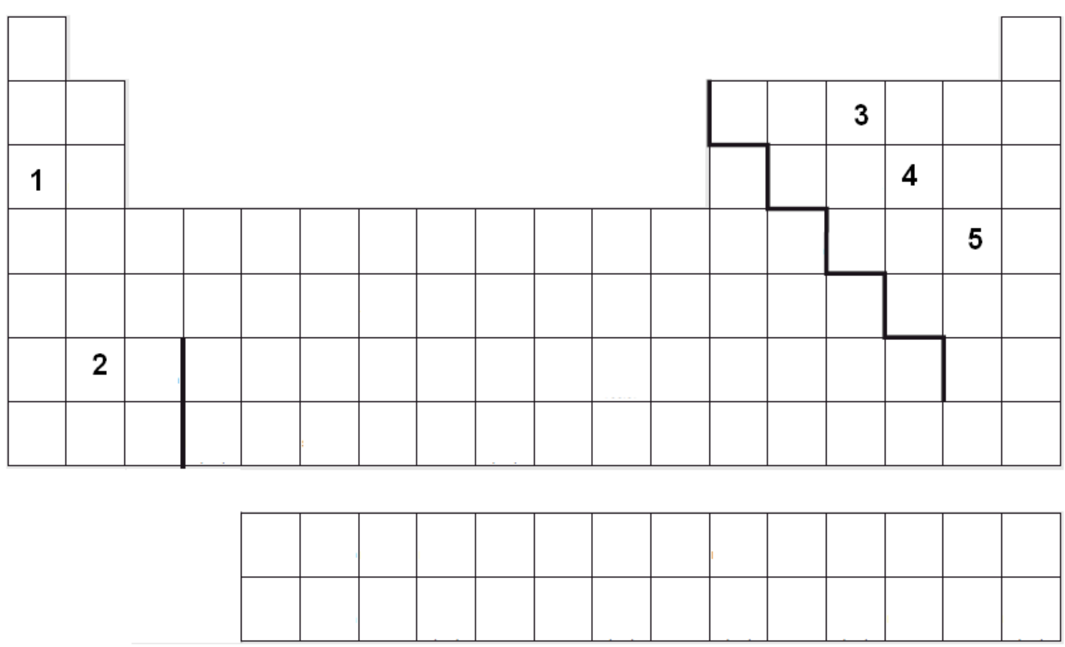
Before Friday's class,

1. Define the following terms:

   orbital

   principal quantum number

2. In the periodic table below: Color the s-block blue, the p-block red, the d-block yellow and the f-block orange. Look carefully at the d-block (transition) metals (fig. 6.30). What is weird about them?



Notes:

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Friday, February 10. Finish Session 3 and Session 4 of GHG Module

Assigned reading: Section 6.5

Quiz today: Nuclear Model of Atom, Bohr Model of the Hydrogen Atom, Nomenclature

Session 3: The second set of concepts that we are addressing today is a combination of the idea of shielding and the idea of orbitals. The concepts of interest are atomic radius, ionization energy and electron affinity. As you look at the chemical equations on p. 279 and p. 282, think about it as a chemical reaction. Identify the reactants and the products. In addition, we are introducing the idea of energy. Ionization reactions always require energy to proceed (indicated by a positive energy value) and while electron affinity reactions could require (+) or produce energy when they proceed (indicated by a negative energy value). We will talk again about sign conventions for energy later in class, but for now concentrate on why it would require energy to remove an electron from an atom.

Session 4: As you read through section 6.5, keep referring back to the periodic table and the information in chapter 2. We are going to look at some reactivity of Alkali and Alkaline Earth metals and the halogens. This chemistry - alkali metals in water - is actually quite exciting (in other words, explosions).

Before Friday's class,

1. Define the following terms:

   Ionization Energy

   Alkali Metal

2. What's the electron configuration of Na? Number of core electrons? Valence (outermost) electrons?

3. Look at the table 6.35 and try to predict which of the alkali metals gives up its 1s electron most easily.

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Problem Set #1

Due Monday, February 13 (at the beginning of class). Late homework will not be accepted.

1. Make a sketch of an s orbital, the p orbitals and the d orbitals.

2. What are the values of the quantum numbers n and l for the following orbitals? **3s, 2p, 3d, 4f**

3. Which set of quantum numbers <u>cannot</u> occur together to specify an orbital?
    a. n=2, l=1, m_l=-1 c. n=3, l=2, m_l=3
    b. n=3, l=3, m_l=2 d. n=4, l=3, m_l=0

4. Which electron is, on average, closer to the nucleus: an electron in a 2s orbital or an electron in a 3s orbital?

5. What is the maximum number of electrons that can have these quantum numbers?
    a. n=3, l=1, ms=+1/2
    b. n=5, l=3
    c. n= 4, m_s = -1/2

6. Write the full electron configuration of the following:
    a. Si c. Rb^+
    b. Mn d. O

7. Draw the full orbital diagrams for the elements with atomic number **10** and **15**. How many unpaired, valence and core electrons does it have?

8. Determine the element that corresponds to each of the following electron configurations:
    a. $[Ar]4s^2\ 3d^{10}\ 4p^6$
    b. $[Kr]5s^2\ 4d^{10}\ 5p^2$

9. Consider these elements: **N, Mg, O, F, Al**.
    a. Write the electron configuration (or atomic orbital diagram) for each element.
    b. Arrange the elements in order of decreasing atomic radius.
    c. Arrange the elements in order of increasing ionization energy.
    d. Use the electron configurations in part a to explain the differences between your answers to part b and c.

10. Potassium is a highly reactive metal while argon is an inert gas. Explain the difference based on their elements configurations.